

 <p><b>IGCSE/GCE</b> <b>The British Programme</b></p>	<p><b>Mayar International Schools</b> <b>First Semester 2025/2026</b></p>	 <p><b>mayar</b> International Schools مدارس ميار الدولية</p>
	<p><b><i>The International Programs'</i></b> <b><i>Department</i></b></p>	

# Atomic Structure and The Periodic Table

## Booklet

**Student Name:** \_\_\_\_\_

**Grade: 11**

## Topic 2: Atomic Structure and the Periodic Table

Students will be assessed on their ability to:

<b>2.1</b>	know the structure of an atom in terms of electrons, protons and neutrons
<b>2.2</b>	know the relative mass and charge of protons, neutrons and electrons
<b>2.3</b>	know what is meant by the terms 'atomic (proton) number' and 'mass number'
<b>2.4</b>	be able to use the atomic number and the mass number to determine the number of each type of subatomic particle in an atom or ion
<b>2.5</b>	understand the term 'isotope'
<b>2.6</b>	<p>understand the basic principles of a mass spectrometer and be able to analyse and interpret mass spectra to:</p> <ul style="list-style-type: none"> <li>i deduce the isotopic composition of a sample of an element</li> <li>ii calculate the relative atomic mass of an element from relative abundances of isotopes and vice versa</li> <li>iii determine the relative molecular mass of a molecule, and hence identify molecules in a sample</li> <li>iv understand that ions in a mass spectrometer may have a 2+ charge</li> </ul>
<b>2.7</b>	be able to predict mass spectra, including relative peak heights, for diatomic molecules, including chlorine, given the isotopic abundances
<b>2.8</b>	be able to define first, second and third ionisation energies and understand that all ionisation energies are endothermic
<b>2.9</b>	know that an orbital is a region within an atom that can hold up to two electrons with opposite spins
<b>2.10</b>	understand how ionisation energies are influenced by the number of protons in the nucleus, the electron shielding and the sub-shell from which the electron is removed
<b>2.11</b>	<p>know that ideas about electronic structure developed from:</p> <ul style="list-style-type: none"> <li>i an understanding that successive ionisation energies provide evidence for the existence of quantum shells and the group to which the element belongs</li> <li>ii an understanding that the first ionisation energy of successive elements provides evidence for electron sub-shells</li> </ul>
<b>2.12</b>	be able to describe the shapes of <i>s</i> and <i>p</i> orbitals
<b>2.13</b>	<p>know that orbitals in sub-shells:</p> <ul style="list-style-type: none"> <li>i each take a single electron before pairing up</li> <li>ii pair up with two electrons of opposite spin</li> </ul>
<b>2.14</b>	be able to predict the electronic configuration of atoms of the elements from hydrogen to krypton inclusive and their ions, using <i>s</i> , <i>p</i> , <i>d</i> notation and electron-in-boxes notation
<b>2.15</b>	understand that electronic configuration determines the chemical properties of an element
<b>2.16</b>	know that the Periodic Table is divided into blocks, such as <i>s</i> , <i>p</i> and <i>d</i> , and know the number of electrons that can occupy <i>s</i> , <i>p</i> and <i>d</i> sub-shells in the first four quantum shells

<b>2.17</b>	be able to represent data, in a graphical form (including the use of logarithms of first ionisation energies on a graph) for elements 1 to 36 and hence explain the meaning of the term 'periodic property'
<b>2.18</b>	<p>be able to explain:</p> <ul style="list-style-type: none"> <li>i the trends in melting and boiling temperatures of the elements of Periods 2 and 3 of the Periodic Table in terms of the structure of the element and the bonding between its atoms or molecules</li> <li>ii the general increase and the specific trends in ionisation energy of the elements across Periods 2 and 3 of the Periodic Table</li> <li>iii the decrease in first ionisation energy down a group</li> </ul>

## Atomic structure and periodic table

### Definitions:

- **The Mass number (A) of an atom** is the given by the sum of the number of protons (Z) and the number of neutrons in the nucleus of an atom.
- **The atomic number (Z)** is given by the number of protons in the nucleus of an atom.
- **Isotopes are atoms of the same element that have the same number of protons but different number of neutrons;** therefore, they have different mass numbers.
- **The relative atomic mass (RAM) is the weighted averages mass of an atom of an element relative to 1/12 the mass of an atom of carbon-12 that has a mass of 12 grams.**
- **Relative isotopic mass is the mass of an individual atom of a particular isotope relative to 1/12 of the mass of an atom of carbon-12.**
- **The relative formula mass (for compounds) is the sum of the relative atomic masses of all the atoms in a chemical formula compared to 1/12 the mass of an atom of carbon-12.**
- **For a molecular compound this is referred to as the relative molecular mass.**

**Note: isotopes of the same elements have identical chemical properties because they have identical electronic configurations.**

## Who discovered electrons, protons and neutrons?

Our current understanding of the structure of atoms is influenced by the theories put forward by scientists such as J.J. Thomson, Ernest Rutherford and James Chadwick.

J.J. Thomson discovered the electron in 1897; Ernest Rutherford discovered the proton in 1917; James Chadwick discovered the neutron in 1932.

## Structure of an atom

Although many other subatomic particles such as quarks, leptons and bosons have since been discovered, chemistry is concerned solely with electrons, protons and neutrons.

The structure of the atom in terms of these three subatomic particles can be summarised as follows.

Particle	Symbol	Relative mass	Relative charge	Position in the atom
proton	p	1	+1	nucleus
neutron	n	1	0	nucleus
electron	e <sup>-</sup>	$\frac{1}{1840}$	-1	energy levels surrounding the nucleus

**table A** The structure of the atom in terms of protons, neutrons and electrons

As well as discovering the proton, Rutherford is credited for first suggesting that the atom has a very small core containing the bulk of the mass of the atom. This core is called the nucleus and contains all of the protons and neutrons that an atom has.

The energy levels surrounding the nucleus in which electrons exist are called quantum shells. We will develop this concept later in this chapter.

All atoms of a given element have the same atomic number, which is different from the atomic number of any other element.

The number of electrons in a neutral atom of an element is equal to the number of protons in the nucleus of that atom.

The electrons surround the nucleus in well-defined energy levels called quantum shells.

1 Use the information in the table below to answer the following questions about particles A to F.

- Which two particles are isotopes of the same element?
- Which two particles are positive ions?
- Which two particles are negative ions?
- Which two particles have the same mass number?

Particle	Number of protons	Number of neutrons	Number of electrons
A	12	13	12
B	17	18	18
C	11	14	10
D	12	12	12
E	35	44	36
F	19	21	18

2 Complete the table to show the numbers of protons, neutrons and electrons in these atoms, molecules and ions.

Symbol for atom or ion	Number of protons	Number of neutrons	Number of electrons
${}^3_1\text{H}$			
${}^{18}_8\text{O}^{2-}$			
${}^{24}_{12}\text{Mg}^{2+}$			
${}^{14}_7\text{N}_1\text{H}_3$			
${}^{14}_7\text{N}_2\text{H}_4^+$			

## Measuring the mass of an atom using mass spectrometer:

To find the relative atomic mass of an element you need to measure its mass and compare it to carbon -12, this is done by using an instrument called **the mass spectrometer**.

### What is a mass spectrometer?

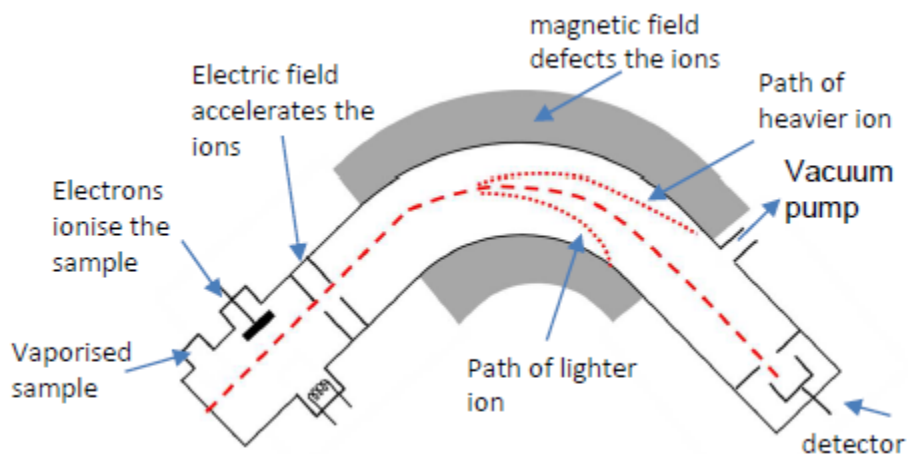
A mass spectrometer measures the masses of atoms and molecules.

It produces positive ions that are deflected by a magnetic field according to their mass-to-charge ratio ( $m/z$ ). It also calculates the relative abundance of each positive ion and displays this as a percentage.

The positive ions could be positively charged atoms, positively charged molecules or positively charged fragments of molecules.

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### How does a mass spectrometer work?



1. The sample is **vaporized**; the sample must be in the **gaseous state** so that its particles can move through the machine.
2. **Ionization:**
  - a) The vapour is ionized by bombardment with high energy electrons which collides with atoms of the sample.

- b) They knock out one or more electrons out of the atoms to form positive ions
3. **Acceleration:** the positive ions are accelerated by an electric field.
  4. They pass through a **velocity selector**; which makes sure they are all travelling at the same velocity, which means any differences in the effect of the magnetic field in the next section is due to different masses or charges of the ions and not to their speeds.
  5. **Deflection:** the ions are deflected by a uniform magnetic field; the amount of deflection depends on the mass to charge ratio of the ions:
    - a) The mass of the ion; the heavier the ion the less it is deflected.
    - b) The charge of the ions; the greater the charge the more it is deflected.

•The degree of deflection depends on the mass-to-charge ratio,  $m/z$ ; the smaller this ratio the larger the deflection.

Ions with large mass to charge ratio are deflected less than the ions with small mass to charge ratio. If all the ions have the same charge, usually +1, the extent of deflection is inversely proportional to their masses.

6. The deflected ions pass through a narrow slit and are collected on a metallic plate connected to an amplifier. For a given strength of magnetic field, only ions with a certain  $m/z$  ratio pass through the slit and are detected. For example an ion with a mass of 28 and a charge of 1+ will be detected at the same time as an ion with a mass of 56 and a charge of 2+. Both ions have  $m/z$  ratio of 28.
7. The strength of the magnetic field is steadily changed to detect positive ions with other  $m/z$  ratios, so only ions of a particular mass: charge ratios pass through to the detector; any other ions will be deflected too much or too little to pass through.
8. The detector detects how many ions pass through the machine. It shows how many ions of each different mass: charge ratio there are in the sample.

In other words, you can determine the exact values of the relative masses of isotopes from a mass spectrum, together with the relative abundance/percentage abundance of each isotope (shown as peaks).

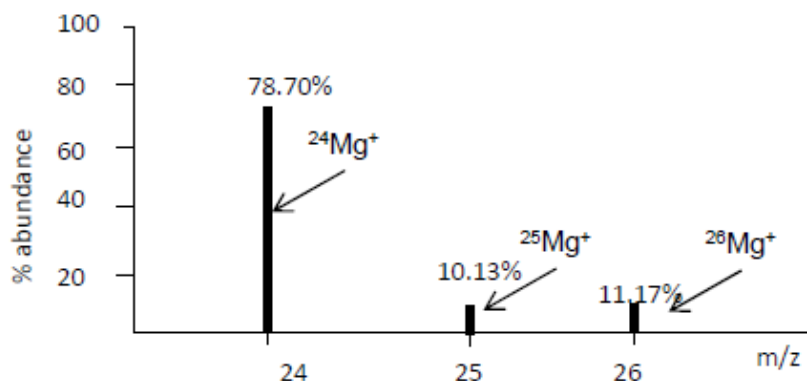


**Question: Explain why there is a vacuum in a mass spectrometer.**

- Because air particles will interfere with the movement of the **ions**/collide with the **ions**/deflect **ions** OR Additional peaks will be detected/peaks at incorrect m/e

### Calculating relative atomic mass

The relative atomic mass quoted on the periodic table is a **weighted average of all the isotopes**



$$\text{R.A.M} = \frac{\sum (\text{isotopic mass} \times \% \text{ abundance})}{100}$$

For above example of Mg

$$\text{R.A.M} = [(78.7 \times 24) + (10.13 \times 25) + (11.17 \times 26)] / 100 = 24.3$$

$$\text{R.A.M} = \frac{\sum (\text{isotopic mass} \times \text{relative abundance})}{\text{total relative abundance}}$$

← If relative abundance is used instead of percentage abundance use this equation

**So:**

**To find the relative atomic masses of elements use this equation:**

$$\text{R.A.M} = \frac{\sum (\text{isotopic mass} \times \% \text{ abundance})}{100}$$

### Activity

1. A sample of iron was analyzed in a mass spectrometer. The table below shows the isotopes it contains. Find the relative atomic mass of this sample.

	Iron-54	Iron-56	Iron-57	Iron-58
Relative abundance (%)	5.84	91.68	2.17	0.13

**Solution:**

1. A sample of copper contains two isotopes of relative isotopic mass 63.0 and 65.0. If the relative atomic mass of copper is 63.5, calculate the relative abundance of each isotope.

### Learning tip

Remember that:

- the mass number and relative isotopic mass are not the same
- the mass number is always a whole number since it is the sum of two whole numbers (the number of protons and the number of neutrons in the nucleus of an atom)
- relative isotopic mass is relative to the mass of a carbon-12 atom and is not likely to be a whole number
- relative isotopic masses should be used to calculate relative atomic masses, but sometimes mass numbers are used to make the arithmetic easier.

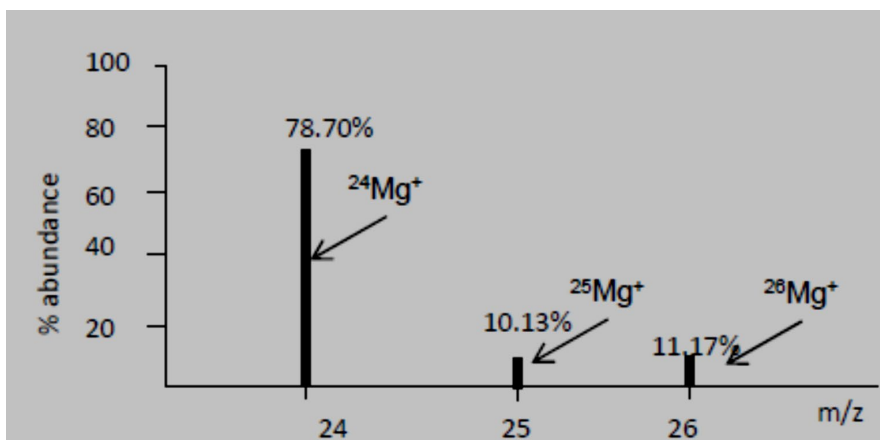
## Finding the relative atomic mass from a mass spectrum:

### THE MASS SPECTRA OF ELEMENTS

#### A. The mass spectrum of monatomic elements

Monatomic elements include all those except for things like chlorine,  $\text{Cl}_2$ , with molecules containing more than one atom.

#### The mass spectrum for magnesium



The 3 peaks in the mass spectrum show that there are 3 isotopes of magnesium with relative isotopic masses of 24, 25 and 26 on the  $^{12}\text{C}$  scale.

For each isotope the mass spectrometer can measure **m/z (mass/charge ratio)** and **abundance**.

**Note:**

- If you are asked to give the species for a peak in a mass spectrum then give **charge** and **mass number** e.g.  $^{24}\text{Mg}^+$
- Sometimes two electrons may be removed from a particle forming a **2+ ion**.  $^{24}\text{Mg}^{2+}$  with a **2+** charge would have **m/z of 12**

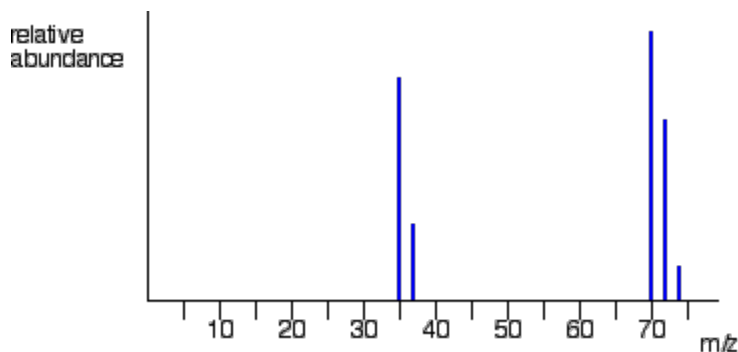
The relative atomic mass of magnesium =  $[(78.7 \times 24) + (10.13 \times 25) + (11.17 \times 26)] / 100 = 24$ .

**B. The mass spectrum of diatomic molecules such as chlorine**

Chlorine is taken as typical of elements with more than one atom per molecule.

Chlorine has two isotopes,  $^{35}\text{Cl}$  and  $^{37}\text{Cl}$  in the approximate ratio of **3 atoms of  $^{35}\text{Cl}$  to 1 atom of  $^{37}\text{Cl}$** . but chlorine consists of molecules, not individual atoms. When chlorine is passed into the ionization chamber, an electron is knocked off the molecule to give a **molecular ion**,  $\text{Cl}_2^+$ . These ions won't be particularly stable, and some will break down into **fragments ( $\text{Cl}^+$  ions)**

The  $\text{Cl}^+$  ions will pass through the machine and will **give lines at 35 and 37**, depending on the isotope; you will also record lines for the **un-fragmented  $\text{Cl}_2^+$  (molecular ions)**.



The **mass spectrum of chlorine** contains **two regions**

- One corresponding to the **Cl atoms** with two lines ( $^{35}\text{Cl}^+$  and  $^{37}\text{Cl}^+$ ).

There are two peaks corresponding to isotopic masses of 35 and 37. The approximate relative peak heights for each isotope are 3:1 for chlorine-35 to chlorine-37.

This gives an approximate relative atomic mass of this sample of chlorine of  $((3 \times 35) + (1 \times 37)) \div 4 = 35.5$ .

- And the other corresponding to the **Cl<sub>2</sub> molecule** with three lines.

Now let's look at the possible combinations of chlorine-35 and chlorine-37 atoms in a Cl<sub>2</sub><sup>+</sup> ion.

Both atoms in the Cl<sub>2</sub><sup>+</sup> molecule could be two of <sup>35</sup>Cl, or two of <sup>37</sup>Cl, or you could have one of each sort (<sup>35</sup>Cl and <sup>37</sup>Cl). That would give you total masses of the Cl<sub>2</sub><sup>+</sup> ion of:

$$35 + 35 = 70 \quad ({}^{35}\text{Cl}-{}^{35}\text{Cl})^+$$

$$35 + 37 = 72 \quad ({}^{35}\text{Cl}-{}^{37}\text{Cl})^+$$

$$37 + 37 = 74 \quad ({}^{37}\text{Cl}-{}^{37}\text{Cl})^+$$

The relative heights of the 70, 72 and 74 lines are in the ratio 9:6:1.

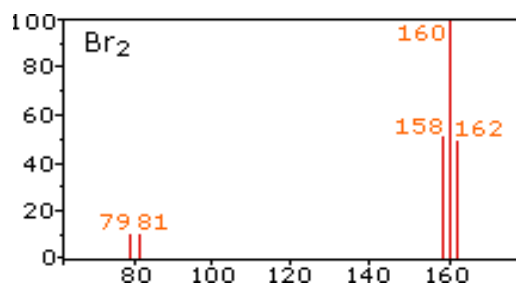
Let's make sense of this.

- There is a 3 in 4 chance of selecting a <sup>35</sup>Cl atom from a sample of chlorine atoms. So the total chance of two <sup>35</sup>Cl atoms combining together is  $\frac{3}{4} \times \frac{3}{4} = \frac{9}{16}$ .
  - There is a 1 in 4 chance of selecting a <sup>37</sup>Cl atom. So the chance of <sup>35</sup>Cl combining with a <sup>37</sup>Cl is  $\frac{1}{4} \times \frac{3}{4} = \frac{3}{16}$ . The chance of a <sup>37</sup>Cl atom combining with a <sup>35</sup>Cl atom is also  $\frac{3}{16}$ . So the total chance of <sup>35</sup>Cl and <sup>37</sup>Cl combining together in any order is  $2 \times \frac{3}{16}$  or  $\frac{6}{16}$ .
  - The chance of two <sup>37</sup>Cl atoms combining together is  $\frac{1}{4} \times \frac{1}{4} = \frac{1}{16}$ .
- $\frac{9}{16} : \frac{6}{16} : \frac{1}{16}$  in whole number ratios is 9:6:1

This corresponds approximately to the peak heights of 70, 72 and 74.

## Mass spectrum of bromine:

Bromine consists of a nearly **50:50** mixture of isotopes having atomic masses of **79** and **81** respectively. Thus, the bromine molecule may be composed of **two  $^{79}\text{Br}$  atoms (mass 158)**, **two  $^{81}\text{Br}$  atoms (mass 162)** or the **more probable combination of  $^{79}\text{Br}$ - $^{81}\text{Br}$  (mass 160)**. Fragmentation of  $\text{Br}_2$  to a bromine cation then gives rise to **equal sized ion peaks at 79 and 81**.



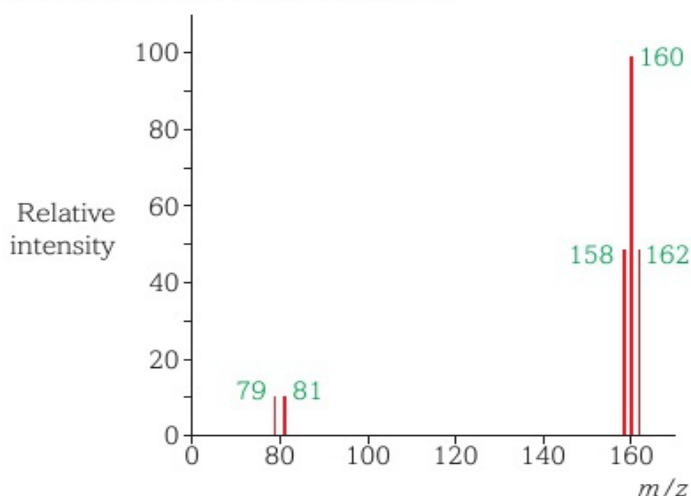
The height at 160 is double the height of 158 and 162.

The heights at 158 and 162 are almost the same.

The heights of 79 and 81 are the same (ratio 50:50)

Question:

The mass spectrum of bromine vapour,  $\text{Br}_2$ , is shown below:



**3 D** Mass spectrum of bromine vapour.

- What are the relative isotopic masses of the two isotopes present in bromine?
- Identify the particles responsible for the peaks at  $m/z$  158, 160 and 162.
- Deduce the relative abundance of the two isotopes and explain the relative heights of the three peaks at  $m/z$  158, 160 and 162.

## The mass spectra of compounds:

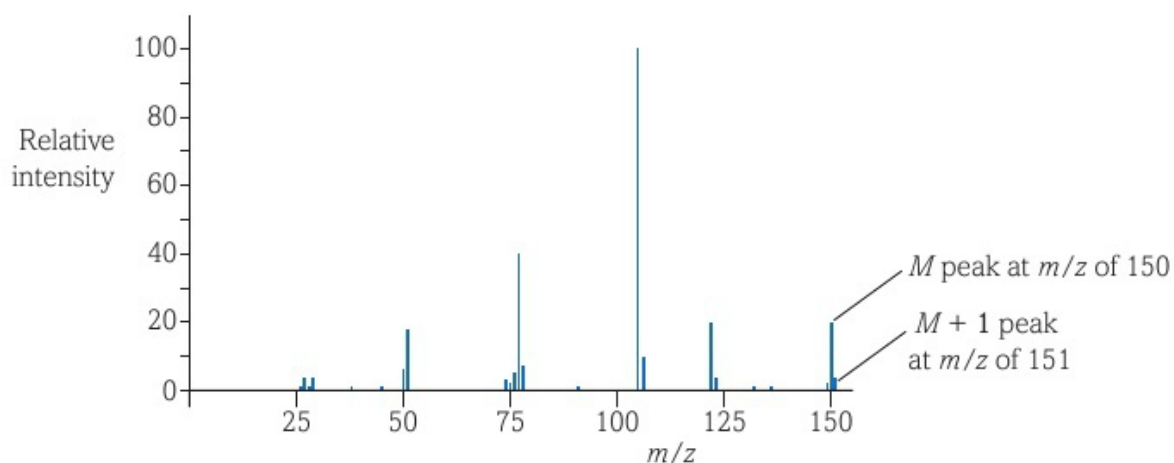
The mass spectrometer can also be used to measure the relative molecular/formula mass of a compound.

When the vaporized sample is bombarded with high- energy electrons, not only is **electrons knocked off**, but the **molecule is fragmented**, leading to a number of peaks corresponding to the various fragments with different masses.

The peak with the largest  $M/Z$  value is due to the **molecular ion,  $M^+$** , (also called the parent ion); it is formed when a sample **molecule loses just electron and is not fragmented**.

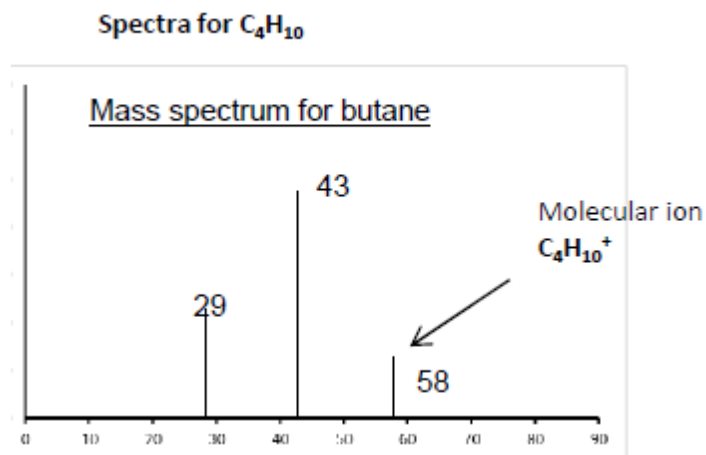
The mass of the parent ion is the  $M_r$  of the compound.

You have to be careful when analysing organic compounds. This is because there is always a small percentage of the carbon-13 isotope present in the compound, which can lead to what is often referred to as an  $M + 1$  peak. This peak can often be seen in molecules with large masses, where the percentage of carbon-13 becomes significant. The peak is often missing, or insignificant, in molecules of small mass.



**fig C** Mass spectrum of ethyl benzoate ( $M_r = 150$ ).

**For example: the mass spectrum of butane shows 3 peaks, the one with the largest  $M/Z$  value is the parent ion ( $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3^+$ ) which has a mass of 58, the other lower  $M/Z$  value correspond to the fragments;  $\text{CH}_3\text{CH}_2\text{CH}_2^+$  and  $\text{CH}_3\text{CH}_2^+$ , with corresponding masses 43 and 29 respectively.**





## Quantum mechanics

**Quantum mechanics** says that the electrons in atoms **are arranged in a series of shells**. Each shell is described by a number known as the **principal quantum number** ( $n=1, n=2, n=3, \dots$ ). These indicate the period number in the periodic table.

**Quantum mechanics** also tells that the principal quantum shells, **apart from the first**, are **split into subshells** which are distinguished by the **letters s, p, d and f**.

**In any principal quantum shell, the energy of the electrons in the subshell increases in the order  $s < p < d$**

**The maximum number of electrons that are allowed in each sub shell is:  $s=2$  electrons,  $p=6$  electrons,  $d=10$  electrons,  $f=14$  electrons.**

**For example:**

**The first quantum shell,  $n=1$ , has only s subshell and can hold a maximum of 2 electrons.**

**Quantum shell,  $n=2$ , has 2 subshells (2s and 2p) and can hold a maximum of 8 electrons (2 electrons in the s subshell + 6 electrons in the p subshell).**

**Quantum shell,  $n=3$ , has 3 sub shells (3s +3p +3d) and can hold a maximum electrons of 18 electrons (2 electrons in the s subshell + 6 electrons in the p subshell + 10 electrons of d subshell)**

The principal quantum shell	Subshells	Maximum number of electrons
$n = 1$	1s	2 electrons
$n = 2$	2s,2p	2+6= 8 electrons
$n = 3$	3s,3p,3d	2+6+10 = 18 electrons
$n = 4$	4s,4p,4d,4f	2+6+10+14= 32 electrons

### Atomic orbitals:

The subshell are further divided into atomic orbitals.

**Definition:**

**Orbital is a region of space around the nucleus where most of the electrons are likely to be found.**

- As each orbital can only hold a maximum number of two electrons, the number of orbitals in each subshell must be:

**s- One orbital**



**p- Three orbitals**



**d- Five orbitals**



**f- Seven orbitals**



**Each orbital is indicated by a single square box.**

- Each orbital is defined by **its energy level, shape and direction in space.**

## Summary:

### Quantum shells

Max Planck first put forward the quantum theory, which is used to describe the arrangement of electrons around the nuclei of atoms.

According to this theory, electrons can only exist in certain well-defined energy levels called **quantum shells**.

All electrons in a given quantum shell have similar, but not identical, energies.

### Electrons in the first four quantum shells

Electrons in the first quantum shell have the lowest energy for that element. The first quantum shell is located in the region closest to the nucleus. The second quantum shell is in a region outside the first; the third is outside the second, and so on.

Each quantum shell, apart from the first, is further divided into subshells of slightly different energy levels.

- There is only one subshell in the first quantum shell. This is labelled the 1s subshell.
- The second quantum shell has two subshells, labelled 2s and 2p. Electrons in the 2p subshell have a slightly higher energy level than those in the 2s subshell.
- The third quantum shell is divided into three subshells, labelled 3s, 3p and 3d. Electrons in the 3p subshell have slightly higher energy than those in 3s, and those in the 3d subshell have slightly higher energy than those in 3p.
- The fourth quantum shell is divided into four subshells, labelled 4s, 4p, 4d and 4f. Once again, the electron energies increase in the order  $4s < 4p < 4d < 4f$ .

## Shapes of the orbitals:

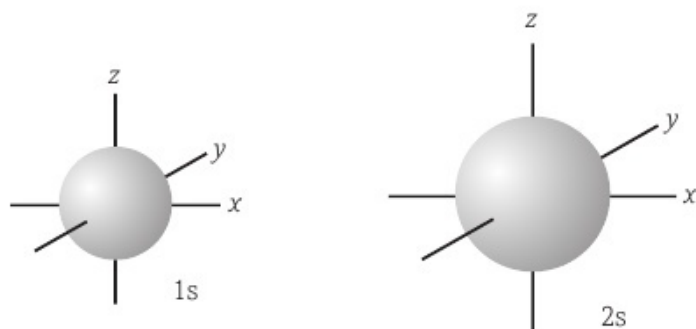
**Each orbital has a three- dimensional shape. Within this shape there is a high probability of finding the electron or the electrons in the orbital.**

### S orbital:

**An s orbital has a spherical shape. The 2s orbital in the second quantum shell has the same shape as the 1s orbital in the 1<sup>st</sup> quantum shell but it is larger.**

**They are both spherical, but the electrons in the 2s orbital have more energy than the electrons in the 1s orbital.**

The sizes and shapes of the orbitals are such that there is a 90% probability of finding the electron within their boundaries.



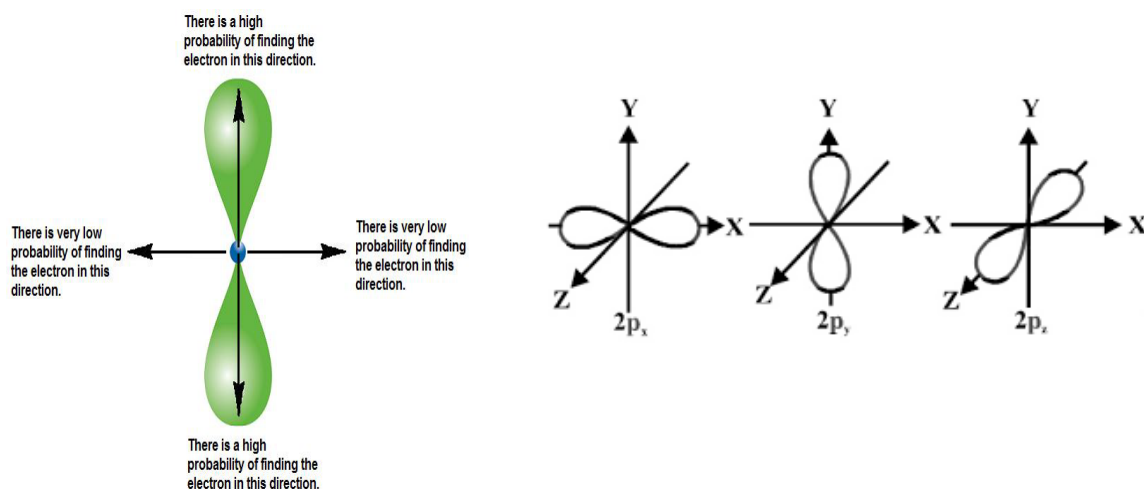
**fig B** The shape of typical 1s and 2s orbitals.

The distribution of electron density in these orbitals **is similar** (they are **spherically symmetrical** around the nucleus).

## P orbitals:

The 2p sub-shell contains three separate p orbital,  $p_x$ ,  $p_y$  and  $p_z$

The p orbitals are approximately dumb-bell shaped, and lie at 90° to each other.



As with the s orbitals, the size of the p orbitals increases with the principal quantum number, so 3p orbital is larger than the 2p orbital.

There can be up to two electrons in each orbital.

# The electronic configuration

Electronic configuration of an atom is the number of electrons in each subshell in each energy level of the atom.

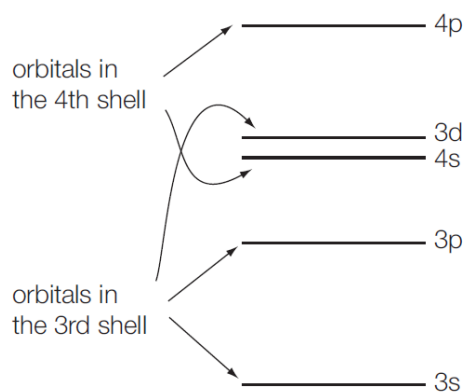
## Filling the subshells and the orbitals:

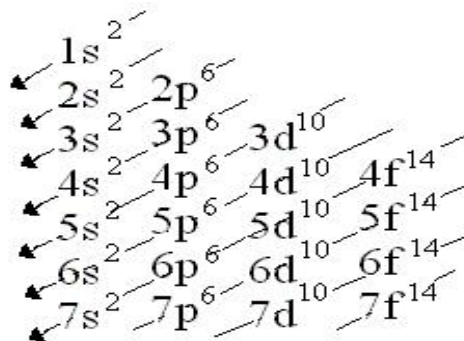
The order in which the subshells are filled depends on their relative energy.

The subshell with the lowest energy, 1s, is therefore filled first (Aufbau principle).

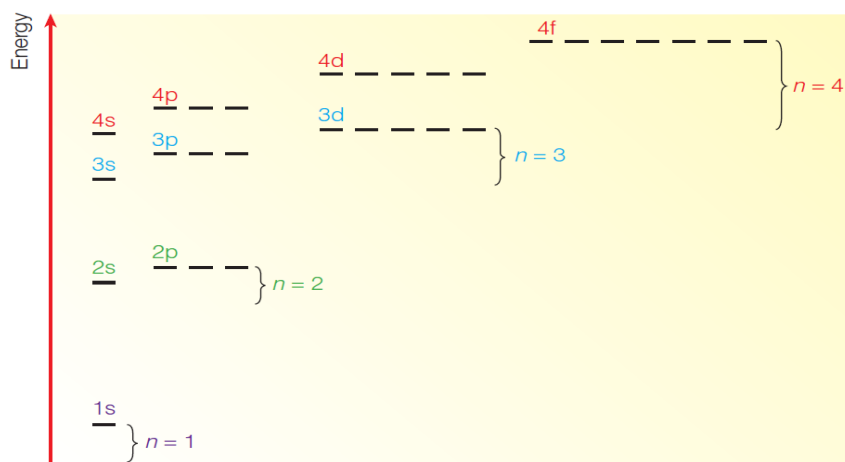
As the shells of electrons around the nuclei of atoms get further from the nucleus, they become closer in energy. Therefore, the difference in energy between the second and third shells is less than that between the first and second.

When the fourth shell is reached there is, in fact, an overlap between the orbitals of the highest energy in the third shell (the 3d orbital) and the lowest energy in the fourth shell (the 4s orbital). In other words, 3d orbitals have an energy level above that of the 4s orbital but below that of 4p orbitals. As a result the orbitals that fill in the **fourth period** are the **4s, 3d and 4p orbitals in that order**.

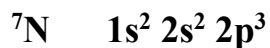
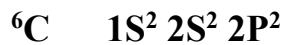
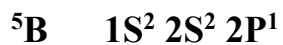
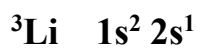
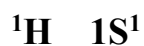




**1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p, 6f, 7d, 7f**



**For example:**



**Electrons are added one by one for successive elements, filling each subshell in order of increasing energy.**

**Exercise:** use the  $1s^2$  notation to write the electronic configuration of the following atoms.

a) vanadium ( $z=23$ )

b) Mg ( $z=12$ )

c) Cl ( $z=17$ )

d) Potassium ( $z=19$ )

a) Zinc ( $z=30$ )

### Filling the orbitals (box form):

A useful way to represent electron configuration is a diagram that places electrons in boxes:

- Each box represents an atomic orbital.
- The boxes (orbitals) can be arranged in order of increasing energy from bottom to top.
- An electron is represented by an arrow,
- The direction of the arrow represents the spin of the electron. We imagine an electron rotating around its axis clockwise or anticlockwise. Pointing up (  $\uparrow$  ) represents spin in one direction and down (  $\downarrow$  ) represents spin in the opposite direction.
- Two electrons in the same orbital cannot have the same spin; this means the orbital can contain only two electrons, having opposite spin.

Chemists have found that paired electrons can only be stable when they spin in opposite directions so that the magnetic attraction resulting from their opposite spins can counteract the electrical repulsion from their negative charges.

**Definition:**

**Hund's rule**, states that the electrons will occupy the orbitals singly before pairing takes place

**The Pauli Exclusion Principle** states that the two electrons cannot occupy the same orbitals unless they have opposite spins

**Write the electron configuration of the following atoms in the box form.**

**Boron ( $z=5$ ), carbon ( $z=6$ ), nitrogen ( $z=7$ ), neon ( $z=10$ )**

**Boron**

**Carbon**

**Nitrogen**

**Neon**

**Exceptions:**

**Cr ( $z=24$ )**

**Cu ( $z=29$ )**



Chromium and copper each only have one 4s electron in their atoms. The explanation for the irregularities lies in the stability of the half filled-filled and filled subshells.

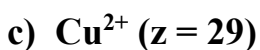
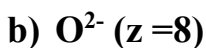
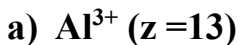
### Writing electron configuration for ions:

- Positive ions are formed when electrons are moved from the atoms. For example, sodium ion,  $\text{Na}^+$  ( $z = 11$ ), has 10 electrons. So, its electron configuration is  $1s^2 2s^2 2p^6$ . Note that this is the same as the electron configuration of neon, the element with 10 electrons in each atom.
- Negative ions are formed when atoms gain electrons. For example, the sulfide ion  $\text{S}^{2-}$  ( $z = 16$ ), has 18 electrons ( $16+2$ ). Its electron configuration is  $1s^2 2s^2 2p^6 3s^2 3p^6$ , which is the same as argon, the element with 18 electrons in each atom.

In general, electrons in the outer shell are removed when metal atoms form their positive ions. However, the d- elements behave slightly differently. Reading across the periodic table from potassium to zinc, the 4s subshells fills before the 3d subshells. But when atoms of a d- block element lose electrons to form ions, the 4s electrons are lost first.

For anions, the electrons in the outer shell are gained when non- metal atoms form their negative ions.

**Exercise: write the electronic configuration for the following ions using the 1s notation:**



d)  $\text{Cu}^+$  ( $z=29$ )

e)  $\text{Cl}^-$  ( $z=17$ )

### Shorthand notation (noble gas notation):

Since chemical changes concern only with the outer electrons in the atom, not the inner shells, so it is helpful to simplify electron configuration to focus on the outer electrons.

For example, the electron configuration of sodium and magnesium are:

Na  $1s^2 2s^2 2p^6 3s^1$

Mg  $1s^2 2s^2 2p^6 3s^2$

The core of electrons in the first two subshells of both atoms is identical- it is only the third subshell that differs.

The inner electrons have the configuration of noble gas neon ( $z=10$ ). The shorthand notation for these elements uses the symbol of the noble gas followed by the outer electrons:

Na:  $[\text{Ne}] 3s^1$

Mg:  $[\text{Ne}] 3s^2$

This method is a shorthand way of writing electronic structures of atoms of many electrons.

**Exercise: write the electron configuration for the following atoms using the shorthand notation:**

Ca ( $z=20$ )

Vanadium ( $z=23$ )

Phosphorous ( $z=15$ )

Beryllium ( $z=4$ )

Arsenic ( $z=33$ )

## Orbitals and the periodic table

Modern periodic law states that the properties of the elements are a function of their atomic number.

The arrangement of elements in the periodic table reflects the electronic structure of the elements.

Groups:

- All elements in a main group (group 1 to group 8) contain the same outer electronic configuration.

	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8
outer electronic configuration	$ns^1$	$ns^2$	$ns^2np^1$	$ns^2np^2$	$ns^2np^3$	$ns^2np^4$	$ns^2np^5$	$ns^2np^6$

**table A** Outer electronic configuration of Groups 1–8 in the Periodic Table.

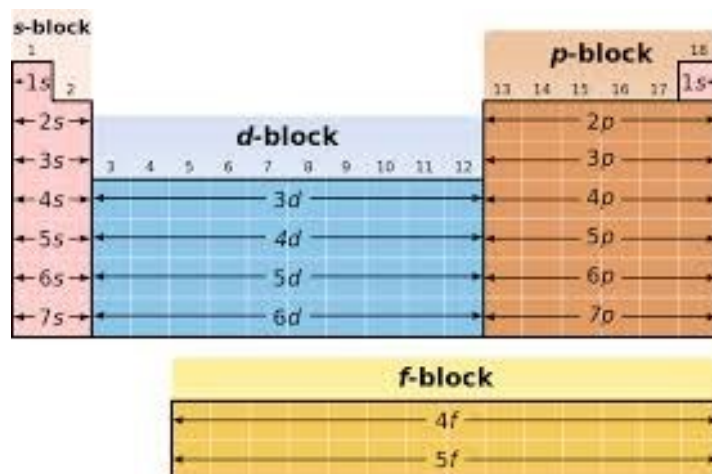
- The elements in the same group have similar chemical properties because they have the same outer electronic structure.

Periods:

- All elements in the same period have the same number of quantum shells containing electrons. For example, all elements in period 2 have electrons in both the first and the second quantum shells.

Blocks:

The periodic table can be divided into blocks, s, p, d and f blocks, determined by orbitals being filled.



## The s- block

The s block contains the elements of group 1 & 2. They are known as the s-block elements because the outer most electrons (the highest energy electrons) are in the s subshell.

These electrons are easily lost to form positive ions.

Hydrogen and helium are also s- block elements; electrons are filling the 1s shells in both. But the characteristics of these two elements- a reactive gas and un- reactive gas, both non-metals are so different from all the other s- block elements.

## The d- block elements

The d- block elements lies between group 2 & 3 and contains the elements with successive electrons being added in d orbitals.

The d- block elements are often called the transition elements; though some of them, such as zinc do not fit this description well.

They are much less reactive than the elements in group 1 & 2.

## **The p – block elements**

The p – block elements are those in group 3, 4, 5, 6, 7, and 8.

- For these elements electrons are being added to the p orbitals in the outer shell.
- Elements in the p block have their highest energy electrons in a p-orbital.

The p – block elements contain all the non – metals (apart from the hydrogen and helium) and metalloids as well as some metals.

However, metals in the p –block do not have strong metallic characteristics, although they all conduct electricity and heat. These post – transition metals are relatively un- reactive and many of their reactions resemble those of non-metals.

## **The f- block elements**

In the f – block elements successive electrons are being added in the f subshells.

They include the lanthanides and the actinides.

## Patterns in the periodic table (periodicity)

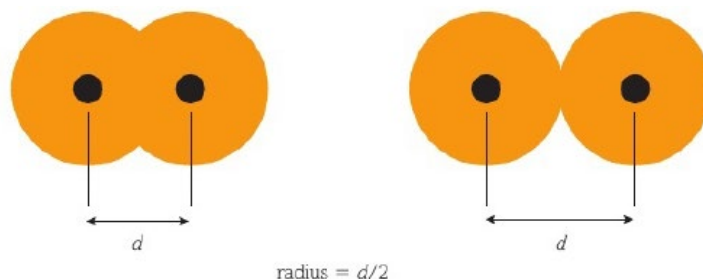
**Periodicity:** regularly repeating trends/patterns of atomic, physical and chemical properties that appear in the elements of the periodic table.

Periodicity can be predicted using the periodic table and explained using the electronic configurations of elements.

## Trends in the periodic table

### Atomic radius:

The atomic radius of an element is a measure of the size of its atoms. It is the distance from the centre of the nucleus to the boundary of the electron cloud. Since the atom does not have a well-defined boundary, we can find the atomic radius by determining the distance between the two nuclei and dividing it by two.



Moving **across the periodic table**, elements **gain electrons**. Moving **down a group**, elements **gain electron shells**. This changes **the size of the atoms**, which in turn affects both their physical and chemical properties.

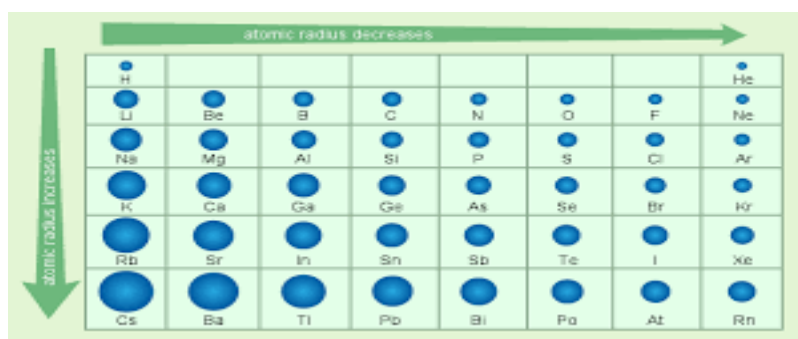
### Trends in the atomic radius of elements:

1. The atomic radius generally decreases across the period (from left to right), this is because

Across the period the nuclear charge becomes increasingly positive, as the number of protons in the nucleus increases. Although the number of electrons also increases, the outer electrons are in the same shell (same shielding by the inner electrons). So they are more strongly attracted to the nucleus, thus reducing the atomic radius.

This increase in the attractive force offsets (counterbalances) the increase in the electron-electron repulsion

2. The atomic radius increases down the group, this is because the outer electrons enter new energy levels (shells) passing down the group. Although the nuclear charge increases down the group, the electrons are further away from the nucleus and being shielded/ screened by more inner electron shells. As a result the outer most electrons are not held tightly to the nucleus, thus the atomic radius increases.



Shielding is an effect of inner electrons which reduces the pull of the nucleus on the electrons in the outer shell of an atom

## Ionic radius:

A positive ion always has a smaller ionic radius than the original atom, this is because:

- The loss of electron/s means that the remaining electrons each have a greater share of the positive charge of the nucleus (same number of protons but less number of electrons), so the outer electrons are more strongly attracted to the nucleus.
- Also, when a positive ion is formed a whole shell is usually lost.

A negative ion has a larger ionic radius than that of the parent atom; this is because

- the addition of extra negative charge (the number of electrons is greater than the number of protons) increases the electron-electron repulsions and so the electrons are less strongly attracted to the nucleus and so the radius increases.

The additional electrons are in the same shell.





## **Ionization:**

**Ionization is the complete removal of an electron from an atom.**

- The electron lost during ionization is so far from the influence of the nucleus (less attracted to the nucleus)
- The energy of the electron has to be increased to a particular value for it to be removed.
- For any given atom, the energy value that the electron has when it reaches this position is always the same, regardless of where in the atom the electron has come from;
  - If the electron already has a high energy (higher quantum shell), then it will not need to gain much energy to be removed.
  - If, however, the electron is in orbital of a low –energy quantum shell, for example the 1s orbital, then it will need to gain a lot of energy to be removed
- The difference in energy between the electron when it has removed and the energy it has when it is in its original ( ground state) orbital in the quantum shell is known as the ionization energy.

### **Definition:**

**Ionization energy is the energy needed to remove 1 mole of electrons from 1 mole of atoms in the gaseous state.**

**This can be measured by gradually increasing the voltage applied to the gas until it conducts and emits light; this means an electron has been removed completely (freed).**

The **First ionization energy** is the energy required to remove **one mole** of **electrons** from **one mole of atoms** in **the gaseous state** to form **+1 ion**.

**For example:**



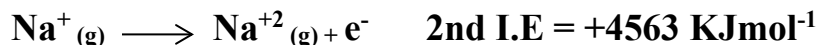
### Things to notice about the equation

The state symbols - (g) - are essential. When you are talking about ionisation energies, everything must be present in the gaseous state.

- **Ionization energies** are measured in  **$\text{kJ mol}^{-1}$**  (kilojoules per mole). They vary in size from **381** (which you would consider very low) up to **2370** (which is very high).
- Ionization is **an endothermic reaction**, since work must be done to overcome the attractive forces between the electron to be removed and the nucleus.

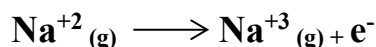
The **second ionization energy** is the energy required to remove 1 mole of electrons from a +1 ion to form +2 ion in the gaseous state.

**For example:**



The total energy required to form a **+2 ion from an atom** is the **sum of the first and second ionization energies**.

The energy required to remove 1 mole of electrons from a +2 ion to form one mole of +3 ion in the gaseous state is called the **third ionization energy**.



We can continue to remove electrons from an atom until only the nucleus is left. We call this sequence of ionization energies, successive ionization energies.

### Factors that affect the ionization energies:

1. The size of the nuclear charge:

- As the atomic number (number of protons) increases, the positive nuclear charge increases. The bigger the positive charge, the greater the attractive force between the nucleus and the outer electrons, leading to higher ionization energy.
- The effect of the increased nuclear charge is to decrease the energy of the electrons since they are strongly attracted to the nucleus.

2. Distance of the outer electrons from the nucleus (the orbital in which the electron exists):

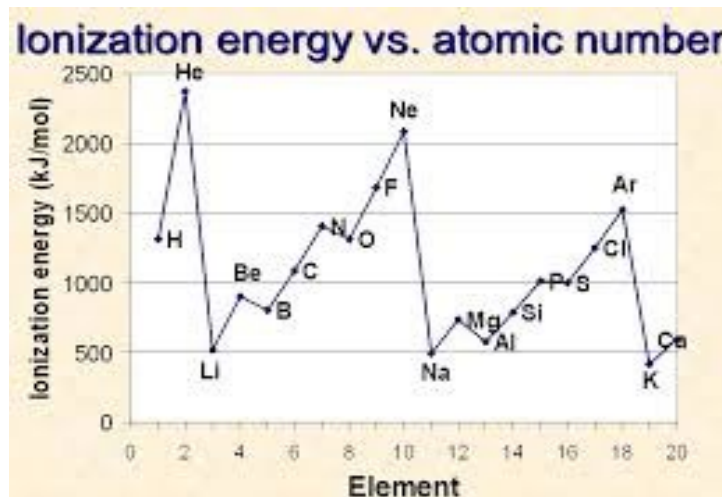
- The force of attraction between the positive nucleus and the outer electrons decreases as the distance between them increases. So, electrons in higher- energy quantum shells are further away from the nucleus and are less attracted to the nucleus than those closer to the nucleus (lower- energy quantum shells).
- The further the outer electron is from the nucleus, the lower the ionization energy.

3. Shielding effect of inner shells (electron-electron repulsion):

- Inner shells electrons **repel** the outer electrons, screening or shielding it from the attraction of the nucleus.
- The electron –electron repulsion shield the other from the effect of the nuclear charge ( the effect of this factor alone would increase the energy of the electrons)
- The more electron shells there are between the outer electrons and the nucleus, the more they experience repulsion/shielding by the inner shell

electrons and though they are less firmly held by the nucleus and the so the first ionization energy decreases.

Trends in ionization energies in the periodic table:



**Patterns across a period:**

1. There is a general increase in the ionization energies across a period, this is the result of:
  - The increasing nuclear charge as the number of protons increases across a period. The atomic radius becomes smaller this leads to an increased attraction between the nucleus and the electron, and therefore a decrease in the energy of the outermost electron and an increase in the ionization energy.
  - The shielding by inner shells; one more electron is added to same quantum shell across the period and this increases the electron-electron repulsion within the quantum shell that would increase the energy in the outermost electron
  - The distance between the outer electrons and the nucleus remains reasonably the same, since the electrons removed come from the same shell.
  - The increase in the nuclear charge is more significant than the electron-electron repulsion, so there is a general increase in the ionization energy across the period.

2. The end of each period is marked by the high ionization energy of a noble gas. This distinctively high ionization energy is a result of the stable electronic structure of the noble gas.
3. There is a rapid decrease in ionization energy between the last element(noble gas) in one period and the first element in the next period; for example the first ionization of lithium is much smaller than helium

He:  $1s^2$

Li:  $1s^2 2s^1$

- The outer electron of lithium is removed from 2s subshell which is in the second quantum shell that is at a higher energy level than the first quantum shell, while as the electron to be removed from the helium is from a lower energy shell, 1s.
- The distance between the nucleus and the outer electron increases.
- The outer most electron of the lithium experience repulsion from the two inner 1s electrons
- These two factors outweigh the increased nuclear charge.

Exception (between group 2&3):

4. There is a slight decrease in the first ionization energy between beryllium and boron (group2 &3), explain why?

Be:  $1s^2 2s^2$

B:  $1s^2 2s^2 2p^1$

- Although the nuclear charge of boron is greater than that of beryllium (has one more proton), the outermost (fifth) electron in boron is removed from 2p subshell, which has more energy than the 2s subshell in the case of beryllium.
- In addition to this, the 2p electron experiences greater electron-electron repulsion (greater shielding by the 1s and 2s electrons), these two factors outweigh the increased nuclear charge.

This is also the case for magnesium and aluminium in period 3 (between group 2 and 3), but instead of 2s and 2p we say 3s and 3p.

Exception (between group 5&6):

5. There is a slight decrease in the ionization energy between nitrogen and oxygen (group 5 &6).

Oxygen has one more proton than nitrogen and the electron removed is from the same 2p subshell. So, you might think that the ionization energy increases, but this is not the case.

O:  $1s^2 2s^2 2p^4$

N:  $1s^2 2s^2 2p^3$

- The electron removed from the nitrogen is from an orbital that contains unpaired electrons.
- The electron removed from the oxygen is one of the two paired electrons in the  $2p_x$  orbital. The extra repulsion between the pair of electrons results in less energy needed to remove an electron. So, the ionization energy for oxygen is lower, because of the spin-pair repulsion.

The presence of two electrons in a single orbital increases the electron-electron repulsion in this orbital so less energy is required to remove one of these electrons.

This is the same case for phosphorus and sulfur in period 3 (between group 5 and six).

In other word, these trends repeat themselves across the third period.

However, the presence of the d- block elements in period 4 disrupts the pattern, as d-block elements have first ionization energies that are relatively the same since the electron to be removed is from the same subshell which is 4s.

You are only responsible for the trends in periods 2 &3

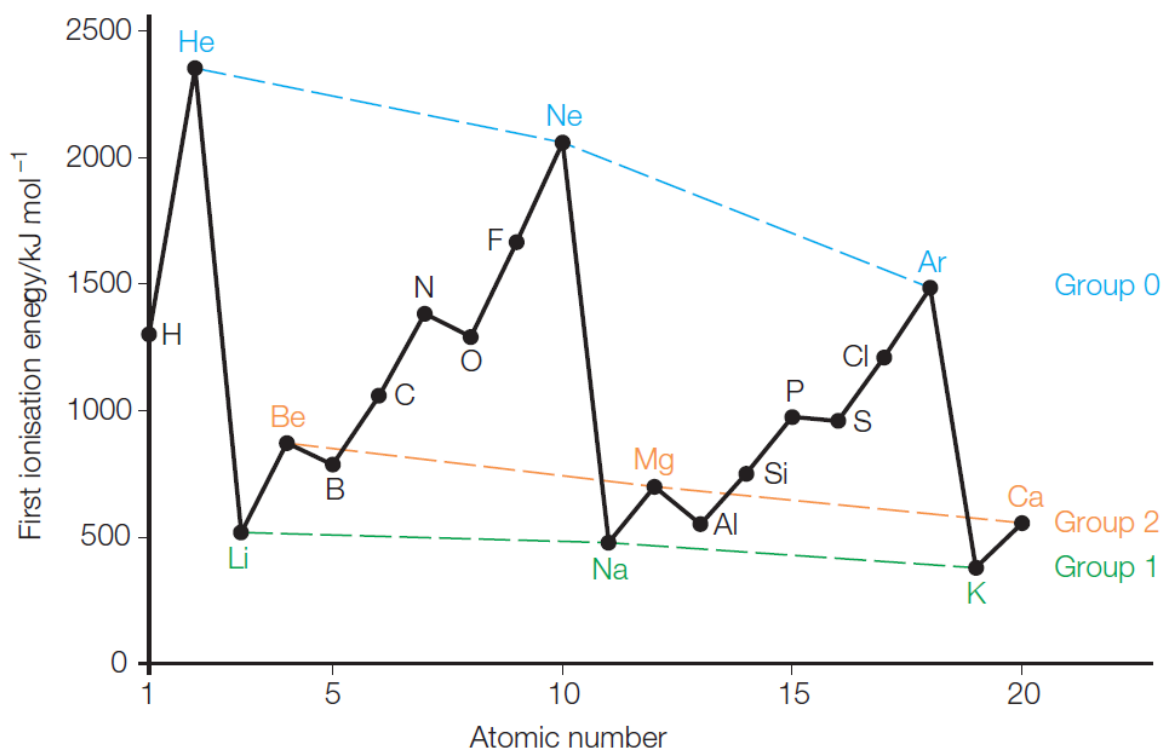
## Patterns down a group:

Although the nuclear charge increases down the group the first ionization energy decreases because:

- Quantum shells are added down the group, this increases the energy of the outermost electrons.
- The outermost electron experiences increased repulsion (increased shielding) from the inner electrons.
- The distance between the nucleus and the outer electrons increases.
- These factors outweigh the increased nuclear charge.

This trend is repeated in group 2, 5, 6, 7 and 8

There is no general trend in first ionization energy in group 3



**Figure 1.29** Periodicity in the first ionisation energies of the elements.

**Q. Why has Helium the largest first ionisation energy?**

A. Its first electron is in the first shell closest to the nucleus and has no shielding effects from inner shells. He has a bigger first ionisation energy than H as it has one more proton

Many questions can be answered by application of the 3 factors that control ionisation energy

**Q. Why do first ionisation energies decrease down a group?**

A. As one goes down a group, the outer electrons are found in shells further from the nucleus and are more shielded so the attraction of the nucleus becomes smaller

**Q. Why is there a general increase in first ionisation energy across a period?**

A. As one goes across a period, the number of protons increases making the effective attraction of the nucleus greater. The electrons are being added to the same shell which has the same shielding effect and the electrons are pulled in closer to the nucleus.

**Q. Why has Na a much lower first ionisation energy than Neon?**

This is because Na will have its outer electron in a 3s shell further from the nucleus and is more shielded. So Na's outer electron is easier to remove and has a lower ionisation energy.

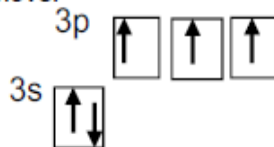
**Q. Why is there a small drop from Mg to Al?**

Al is starting to fill a 3p sub shell, whereas Mg has its outer electrons in the 3s sub shell. The electrons in the 3p subshell are slightly easier to remove because the 3p electrons are higher in energy and are also slightly shielded by the 3s electrons

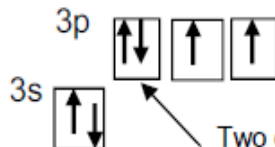
**Q. Why is there a small drop from P to S?**

With sulphur there are 4 electrons in the 3p sub shell and the 4th is starting to doubly fill the first 3p orbital.

When the second electron is added to a **3p orbital** there is a slight **repulsion** between the two negatively charged electrons which makes the second electron easier to remove.



phosphorus 1s<sup>2</sup> 2s<sup>2</sup> 2p<sup>6</sup> 3s<sup>2</sup> 3p<sup>3</sup>



Two electrons of opposite spin in the same orbital

sulphur 1s<sup>2</sup> 2s<sup>2</sup> 2p<sup>6</sup> 3s<sup>2</sup> 3p<sup>4</sup>

Learn carefully the explanations for these two small drops as they are different to the usual factors



## Successive ionization energy:

The different energies required to remove the first and the subsequent electrons from an atom is called the successive ionization energy.

### Ionization Energies (kJ/mol)

Element	1 <sup>st</sup>	2 <sup>nd</sup>	3 <sup>rd</sup>	4 <sup>th</sup>	5 <sup>th</sup>	6 <sup>th</sup>
Na	498	4560	6910	9540	13,400	16,600
Mg	736	1445	7730	10,600	13,600	18,000
Al	577	1815	2740	11,600	15,000	18,310
Si	787	1575	3220	4350	16,100	19,800
P	1063	1890	2905	4950	6270	21,200
S	1000	2260	3375	4565	6950	8490
Cl	1255	2295	3850	5160	6560	9360
Ar	1519	2665	3945	5770	7320	8780

Heron, Frank, Sarquis, Sarquis, Cochran, Kulka, Chemistry, 1996, Heath, page

□ Shaded area on table denotes core electrons.

The successive ionization energies for the elements in period 3:

1. For each element the successive ionization energies increases. This is because:

- the charge on the ions gets greater as each electron is removed (the number of protons is greater than the number of electrons).
- As each electron is removed the electron –electron repulsion within the same quantum shell decreases (less shielding), this results in a decrease in the energies of the remaining electrons and therefore a steady increase in the ionization energy within the same quantum shell.
- There is a greater attractive force between the positively charged protons in the nucleus and the remaining negatively charged electrons (same number of protons but less number of electrons). Therefore, more energy is needed to overcome these attractive forces.

2. There is a big difference (jump) between some successive ionization energies for each element.

For example for sodium there is a big gap (difference) between the first and second ionization energies, these large changes indicate that the second electron is being removed from a shell closer to the nucleus (the electron has a lower energy and experience less repulsion ). The electron configuration of sodium is  $1s^2 2s^2 2p^6 3s^1$

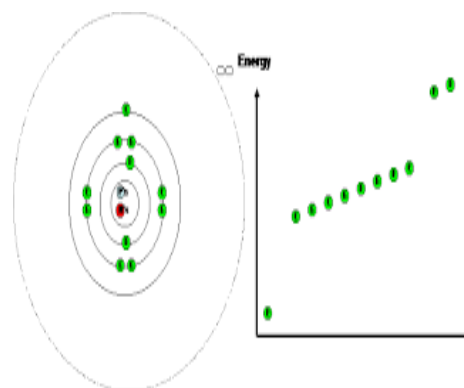
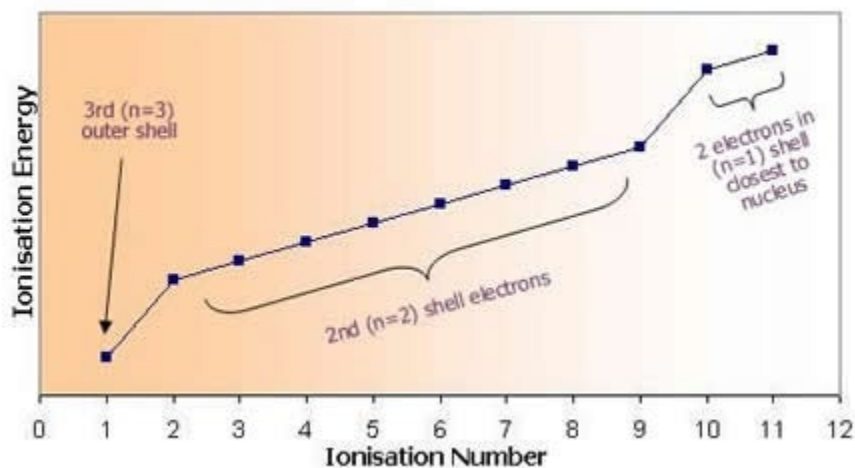
Another example is the big difference (gap) between the fifth and the sixth ionization energy for phosphorus; the fifth electron of phosphorus being removed from the third quantum principal shell, while as the sixth electron being removed from the second quantum shell which is closer to the nucleus.

P:  $1s^2 2s^2 2p^6 3s^2 3p^3$

- You can work out which group an element is in from the first big jump in its successive energies.

The patterns in successive ionization energies for an element give us important information about the electronic structure for that element (shells and the proton number).

For example, here is a graph of the successive ionization energies for sodium against the number of electrons removed.



We can deduce the following about sodium:

- The first electron removed has a lowest ionization energy (compared with other data), which indicates that it is likely to be further away from the nucleus (third quantum shell) and has the highest energy of any electrons in the sodium atom and well shielded by inner electron shells.
- The second electron is much more difficult to remove than the first electron. There is a big jump in the value of the first and the second ionization energies; this suggests the second electron is in a shell closer to the nucleus (the electron is to be removed from a quantum shell with lower energy - second quantum shell) than the first electron and so is not in the same shell as the first one. Taken together the first and second ionization energies suggest that sodium has one electron in its outer shell (group 1).
- From the second to the ninth electrons removed there is only a gradual increase in successive ionization energies, this suggests that all the eight electrons are in the same quantum shell. As each electron is removed the electron –electron repulsion within the same quantum shell decreases (less shielding), this results in a decrease in the energies of the remaining electrons and therefore a steady increase in the ionization energy.
- There is a big jump between the 9th and 10th ionization energies; this suggests that the 10<sup>th</sup> electron is in a quantum shell closer to the nucleus than the 9<sup>th</sup> electron (has lower energy and less shielded).
- The 10<sup>th</sup> and 11<sup>th</sup> electrons have extremely high ionization energies. This suggests that they are very close to the nucleus; there must be a very great force of attraction between the nucleus and these electrons and there are no inner shells to shield them.
- So, the electron configuration of sodium is 2,8,1

Note: the effect of electron-electron repulsion is to raise the energy of the electrons involved above the value they would have if there were no repulsion between them.

We can use successive ionization energies to:

1. Predict or to confirm the simple electronic configuration of elements.
2. Confirm the number of electrons in the outer shell of an element and hence the group number to which it belongs.
3. Successive ionization energies of a particular element provide evidence for the existence of quantum shells.

The first ionization energy of successive elements provides evidence for electron sub-shells.

Important notes:

1. The first ionization energy is a measure of how tightly or loosely the outer electron is attracted to the nucleus. The less tightly it is bound (the more easily it is removed) the more reactive the element will be.
2. The first ionization energy is the lowest because an electron from an outer shell, furthest from the nucleus, is removed first.
3. Electrons closer to the nucleus are more tightly attracted to the nucleus because few electron shells are there between them and the nucleus, so they are less shielded (experience less electron-electron repulsion) from the positive attraction of the protons.
4. The successive ionization energies help to confirm that the electrons are to be found at different energy level (shells).

Questions:

1. The successive ionization energy of boron are shown in the table below:

ionization	1 <sup>st</sup>	2 <sup>nd</sup>	3 <sup>rd</sup>	4 <sup>th</sup>	5 <sup>th</sup>
<b>Ionization energy/KJmol<sup>-1</sup></b>	<b>799</b>	<b>2420</b>	<b>3660</b>	<b>25000</b>	<b>32800</b>

- a) Why is there a large increase between the third and fourth ionization energies:

b) Explain how these figures confirm that the electronic of boron is 2, 3.

2. For the element aluminium (  $z=13$ ), draw a sketch graph to predict the  $\log_{10}$  of the successive ionization energies against the number of electrons removed

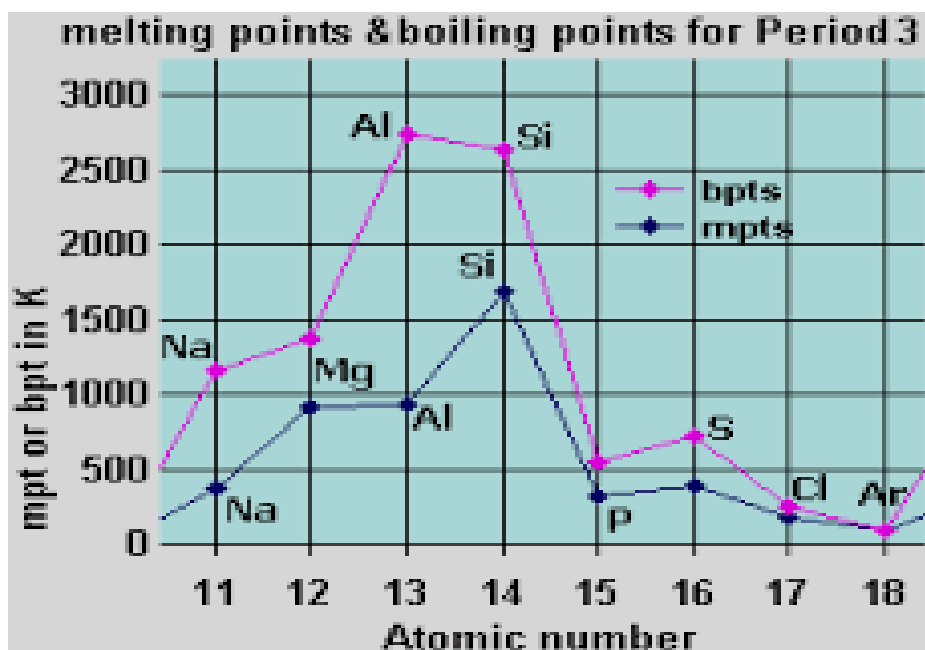
3. The successive ionization energy of an element x are shown in the table below. Which group in the periodic table does x belong to?

ionization	1 <sup>st</sup>	2 <sup>nd</sup>	3 <sup>rd</sup>	4 <sup>th</sup>	5 <sup>th</sup>	6 <sup>th</sup>	7 <sup>th</sup>	8 <sup>th</sup>	9 <sup>th</sup>	10 <sup>th</sup>
<b>Ionization energy</b>	<b>1000</b>	<b>2260</b>	<b>3390</b>	<b>4540</b>	<b>7010</b>	<b>8500</b>	<b>27100</b>	<b>31670</b>	<b>36580</b>	<b>43140</b>

## Trends in the melting points and boiling points (period 2 &3):

Physical properties, such as the melting and boiling points and the electrical conductivity of the elements, also show trends across the period.

Taking period 3 as an example:



The electrical conductivity increases across the metals of period 3 from sodium to aluminium. The electrical conductivity then drops dramatically to silicon, which is described as a semiconductor, and falls even further to the non-metallic insulators phosphorus and sulfur.

To explain the trend in **melting points** and **electrical conductivity across a period**, we have to consider the **bonding** and the **structure** of the elements.

PERIOD 2	Li	Be	B	C (DIAMOND)	N	O	F
melting temperature / °C	181	1278	2300	3550	-210	-218	-220
boiling temperature / °C	1342	2970	3927	4827	-196	-183	-188
type of bonding	metallic	metallic	covalent	covalent	covalent	covalent	covalent
structure	giant lattice	giant lattice	giant lattice	giant lattice	simple molecular	simple molecular	simple molecular

PERIOD 3	Na	Mg	Al	Si	P	S	Cl
melting temperature / °C	98	649	660	1440	44	113	-101
boiling temperature / °C	883	1107	2467	2355	280	445	-35
type of bonding	metallic	metallic	metallic	covalent	covalent	covalent	covalent
structure	giant lattice	giant lattice	giant lattice	giant lattice	simple molecular	simple molecular	simple molecular

- **Sodium, magnesium and aluminium**, at the start of period 3, are all **metallic** elements, which accounts for their **electrical conductivity** and relatively **high melting and boiling points**.
- **Melting and boiling points rise** across the three metals because of the increasing number of electrons which each atom can contribute to the delocalised "sea of electrons". The atoms also get smaller and have more protons as you go from sodium to magnesium to aluminium.

The attractions and therefore the **melting and boiling points increase**, this can be explained by:

- The number of electrons each metal donates into the sea of delocalized electrons.
- And the increasing charge on the metal ions in the giant metallic structure.

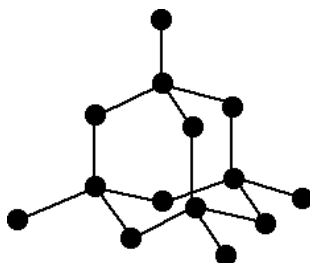
For example, sodium atom donates just 1 electron, forming  $\text{Na}^+$  ions in the lattice, magnesium atom donates 2 electrons, forming  $\text{Mg}^{2+}$  ions, whereas each aluminium atom donates 3 electrons to the sea of delocalized electrons, forming  $\text{Al}^{3+}$  ions. This makes the metallic bonding in aluminium the strongest, as the electrostatic forces of attraction between its  $3^+$  ions and the larger number of negatively charged delocalized electrons holding the giant structure together is stronger.

There are also more delocalized electrons available to drift through the structure when aluminium metal conducts an electric current, making aluminium a better conductor than sodium or magnesium.

- Moving across the periods you meet giant molecular structures like those found in the metalloids silicon and carbon in the form of diamond; they have many strong covalent bonds between the atoms which hold

them tightly within a crystal structure. As a result it is very difficult to remove individual atoms-these elements have extremely high melting temperatures.

- However, their **electrical conductivity** is **much lower** than the metals at the start of the period because there are no delocalized electrons free to move around within their structures.



- In contrast, the simple molecular structures found in most non-metals on the right of the periodic table mean they exist as small, individual molecules.
- Although the forces holding the atoms within the molecules are strong covalent forces, the molecules themselves are only held together by weak inter molecular forces (Vander Waal forces), so less energy is needed to break these bonds and as a result most non-metals have low melting temperatures.

### Four molecular elements

Phosphorus, sulfur, chlorine and argon are **simple molecular** substances with only van der Waals attractions between the molecules. Their **melting or boiling points will be lower than those of the first four members of the period which have giant structures.** The presence of individual molecules prevents any possibility of electrons flowing, and so **none of them conduct electricity.**

The **sizes of the melting and boiling points** are governed entirely by the sizes of the molecules:





a P<sub>4</sub> molecule



an S<sub>8</sub> molecule



a Cl<sub>2</sub> molecule

Argon molecules consist of **single argon** atoms.

## Phosphorus

There are several forms of phosphorus. The data in the graph at the top of the page applies to white phosphorus which contains **P<sub>4</sub> molecules**. **To melt phosphorus you don't have to break any covalent bonds - just the much weaker van der Waals forces between the molecules.**

## Sulphur

**Sulphur consists of S<sub>8</sub> rings of atoms.** The molecules are **bigger** than phosphorus molecules, and so **the van der Waals attractions will be stronger**, leading to a **higher melting and boiling point**.

## Chlorine

**Chlorine, Cl<sub>2</sub>, is a much smaller molecule with comparatively weak van der Waals attractions, and so chlorine will have a lower melting and boiling point than sulphur or phosphorus.**

## Argon

**Argon molecules are just single argon atoms, Ar.** The van der Waals attraction between these is very limited and so the melting and boiling points of argon are **lower again**.

